



OFFICIAL STUDY GUIDE 2001 EDITION



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COLLEGE-LEVEL EXAMINATION PROGRAM

General Chemistry

Description of the Examination

The Subject Examination in General Chemistry covers material usually taught in a one-year introductory course in general chemistry. Understanding of the structure and states of matter, reaction types, equations and stoichiometry, equilibrium, kinetics, thermodynamics, and descriptive and experimental chemistry is required, as is the ability to interpret and apply this material to new and unfamiliar problems.

Battery-operated, hand-held calculators may be used during the exam; however, all calculator memories must be cleared of programs and data, both before and after the exam. It may be helpful to use a calculator for some questions on the exam. A Periodic Table of the elements is included in the exam booklet.

The exam contains approximately 80 multiple-choice questions to be answered within two separately timed 45-minute periods.

Knowledge and Skills Required

Questions on the exam require candidates to demonstrate one or more of the following abilities.

- Recall: remember specific facts; demonstrate straightforward knowledge of information and familiarity with terminology
- Application: understand concepts and reformulate information into other equivalent terms; apply knowledge to unfamiliar and/or practical situations; solve mathematical problems
- Interpretation: infer and deduce from data available and integrate information to form conclusions; recognize unstated assumptions

The subject matter of the General Chemistry exam is drawn from the following topics.

➡	<i>Approximate Percent of Examination</i>
20%	<p>Structure of Matter</p> <ul style="list-style-type: none"> Atomic theory and atomic structure <ul style="list-style-type: none"> Evidence for the atomic theory Atomic masses; determination by chemical and physical means Atomic number and mass number; isotopes Electron energy levels: atomic spectra, quantum numbers, atomic orbitals Periodic relationships including, for example, atomic radii, ionization energies, electron affinities, oxidation states Chemical bonding <ul style="list-style-type: none"> Binding forces <ul style="list-style-type: none"> Forces within species: covalent, ionic, metallic Intermolecular forces: hydrogen bonding, dipole-dipole, London dispersion forces Relationships to states, structure, and properties Polarity of bonds, electronegativities Molecular models <ul style="list-style-type: none"> Lewis structures Valence bond theory; hybridization of orbitals, resonance, sigma and pi bonds VSEPR Geometry of molecules, ions, and coordination complexes; dipole moments of molecules; relation of properties to structure Nuclear chemistry: nuclear equations, half-lives, and radioactivity; chemical applications

➤ *Approximate Percent of Examination*

- 19% States of Matter
- Gases

Laws of ideal gases; equations of state for an ideal gas; partial pressures

Kinetic-molecular theory

Interpretation of ideal gas laws on the basis of this theory

Avogadro's hypothesis and the mole concept

Dependence of kinetic energy of molecules on temperature

Deviations from ideal gas laws
 - Liquids and solids

Liquids and solids from the kinetic-molecular viewpoint

Phase diagrams of one-component systems

Changes of state, including critical points and triple points

Structure of solids; lattice energies
 - Solutions

Types of solutions and factors affecting solubility

Methods of expressing concentration

Raoult's law and colligative properties (nonvolatile solutes); osmosis

Non-ideal behavior (qualitative aspects)
- 38% Reactions
- Reaction types (12%)

Acid-base reactions; concepts of Arrhenius, Brönsted-Lowry, and Lewis; amphoterism; coordination complexes

Precipitation reactions

Oxidation-reduction reactions

The role of the electron in oxidation-reduction

Electrochemistry: electrolytic and voltaic cells; Faraday's laws; standard half-cell potentials; Nernst equation; prediction of the direction of redox reactions

➤ *Approximate Percent of Examination*

- Equations and Stoichiometry (10%)
 - The mole concept; Avogadro's number
 - Ionic and molecular species present in chemical systems; net ionic equations
 - Balancing of equations including those for redox reactions
 - Mass and volume relations with emphasis on the mole concept, including empirical formulas and limiting reactants
- Equilibrium (7%)
 - Concept of dynamic equilibrium, physical and chemical; Le Chatelier's principle; equilibrium constants
 - Quantitative treatment
 - Equilibrium constants for gaseous reactions in terms of both molar concentrations and partial pressure (K_c , K_p)
 - Equilibrium constants for reactions in solutions
 - Constants for acids and bases; pK ; pH
 - Solubility product constants and their application to precipitation and the dissolution of slightly soluble compounds
 - Common ion effect; buffers; hydrolysis
- Kinetics (4%)
 - Concept of rate of reaction
 - Use of differential rate laws to determine order of reaction and rate constant from experimental data
 - Effect of temperature change on rates
 - Energy of activation; the role of catalysts



Approximate Percent of Examination

- Thermodynamics (5%)

State functions

First law: heat of formation; heat of reaction; change in enthalpy; Hess's law; heats of vaporization and fusion

Second law: free energy of formation; free energy of reaction; dependence of change in free energy on enthalpy and entropy changes

Relationship of change in free energy to equilibrium constants and electrode potentials

14% Descriptive Chemistry

- The accumulation of certain specific facts of chemistry is essential to enable students to comprehend the development of principles and concepts, to demonstrate applications of principle, to relate fact to theory and properties to structure, and to develop an understanding of systematic nomenclature, which facilitates communication. The following areas are normally included on the exam.

Chemical reactivity and products of chemical reactions

Chemistry of the main groups and transition elements, including typical examples of each

Organic compounds as exemplary material in areas such as bonding, acid-base reactions, structure, solutions, intermolecular forces

9% Experimental Chemistry

- Some questions are based on laboratory experiments widely performed in general chemistry and ask about the equipment used, observations made, calculations performed, and interpretation of the results. The questions are designed to provide a measure of students' understanding of the basic tools of chemistry and their applications to simple chemical systems.

Sample Questions

The 46 sample questions that follow are similar to questions on the General Chemistry exam, but they do not appear on the actual exam. CLEP exams are designed so that average students completing a course in the subject can usually answer about half the questions correctly. Because a Periodic Table is provided with the actual exam, you should refer to a Periodic Table as needed when working through the sample questions.

Before attempting to answer the sample questions, read all the information about the General Chemistry exam on the preceding pages. Additional suggestions for preparing for CLEP exams are provided in Chapter 1.

Try to answer correctly as many questions as possible. Then compare your answers with the correct answers, given at the end of this examination guide.

Note: For all questions involving solutions and/or chemical equations, assume that the system is in pure water and at room temperature unless otherwise stated.

Part A

Directions: Each set of lettered choices below refers to the numbered questions or statements immediately following it. Select the one lettered choice that best answers each question or best fits each statement. A choice may be used once, more than once, or not at all in each set.

Questions 1-3

- (A) F
- (B) S
- (C) Mg
- (D) Ar
- (E) Mn

1. Forms monatomic ions with 2^- charge in solutions (A) (B) (C) (D) (E)
2. Forms a compound having the formula KXO_4 (A) (B) (C) (D) (E)
3. Forms oxides that are common air pollutants and that yield acidic solutions in water (A) (B) (C) (D) (E)

Questions 4-6

- (A) Hydrofluoric acid
 - (B) Carbon dioxide
 - (C) Aluminum hydroxide
 - (D) Ammonia
 - (E) Hydrogen peroxide
4. Is a good oxidizing agent (A) (B) (C) (D) (E)
5. Is used extensively for the production of fertilizers (A) (B) (C) (D) (E)
6. Has amphoteric properties (A) (B) (C) (D) (E)

Questions 7-8

- (A) A network solid with covalent bonding
 - (B) A molecular solid with zero dipole moment
 - (C) A molecular solid with hydrogen bonding
 - (D) An ionic solid
 - (E) A metallic solid
7. Solid ethyl alcohol, C_2H_5OH (A) (B) (C) (D) (E)
8. Silicon dioxide, SiO_2 (A) (B) (C) (D) (E)

Questions 9-11

- (A) CO_3^{2-}
- (B) MnO_4^-
- (C) NH_4^+
- (D) Ba^{2+}
- (E) Al^{3+}

Assume that you have an “unknown” consisting of an aqueous solution of a salt that contains one of the ions listed above. Which ion must be absent on the basis of each of the following observations of the “unknown”?

9. The solution is colorless. (A) (B) (C) (D) (E)

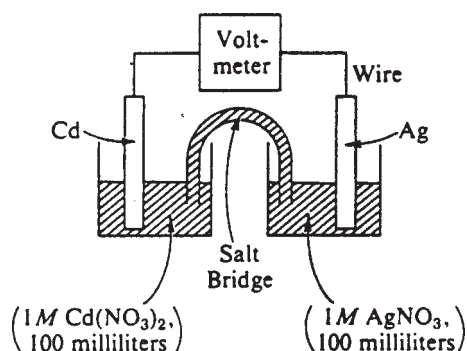
10. No odor can be detected when a sample of the solution is added drop by drop to a warm solution of sodium hydroxide.

(A) (B) (C) (D) (E)

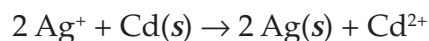
11. No precipitate is formed when a dilute solution of H_2SO_4 is added to a sample of the solution.

(A) (B) (C) (D) (E)

Questions 12-13



The spontaneous reaction that occurs when the cell above operates is



- (A) Voltage increases.
- (B) Voltage decreases but remains above zero.
- (C) Voltage becomes zero and remains at zero.
- (D) No change in voltage occurs.
- (E) Direction of voltage change cannot be predicted without additional information.

Which of the above occurs for each of the following circumstances?

12. The silver electrode is made larger. (A) (B) (C) (D) (E)

13. The salt bridge is replaced by a platinum wire. (A) (B) (C) (D) (E)

Part B

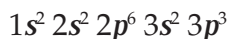
Directions: Each of the questions or incomplete statements below is followed by five suggested answers or completions. Select the one that is best in each case.

14.	Hydrogen Halide	Normal Boiling Point, °C
	HF	+19
	HCl	-85
	HBr	-67
	HI	-35

The liquefied hydrogen halides have the normal boiling points given above. The relatively high boiling point of HF can be correctly explained by which of the following?

- (A) HF gas is more ideal.
 (B) HF is the strongest acid.
 (C) HF molecules have a smaller dipole moment.
 (D) HF is much less soluble in water.
 (E) HF molecules tend to form hydrogen bonds.

(A) (B) (C) (D) (E)



15. Atoms of an element, X, have the electronic configuration shown above. The compound most likely formed with magnesium, Mg, is

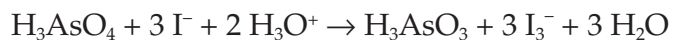
- (A) MgX (B) Mg₂X (C) MgX₂
 (D) Mg₂X₃ (E) Mg₃X₂

(A) (B) (C) (D) (E)

16. The density of an unknown gas is 4.20 grams per liter at 3.00 atmospheres pressure and 127°C. What is the molecular mass of this gas? ($R = 0.0821$ liter-atm/mole-K)

- (A) 14.6
 (B) 46.0
 (C) 88.0
 (D) 94.1
 (E) 138

(A) (B) (C) (D) (E)

Questions 17-18

The oxidation of iodide ions by arsenic acid in acidic aqueous solution occurs according to the stoichiometry shown above. The experimental rate law for the reaction is:

$$\text{Rate} = k[\text{H}_3\text{AsO}_4] [\text{I}^-] [\text{H}_3\text{O}^+]$$

17. What is the order of the reaction with respect to I^- ?

- (A) 1 (B) 2 (C) 3 (D) 5 (E) 6 (A) (B) (C) (D) (E)

18. According to the rate law for the reaction shown above, an increase in the concentration of hydronium ion has what effect on this reaction?

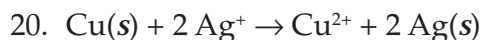
- (A) The rate of reaction increases.
(B) The rate of reaction decreases.
(C) The value of the equilibrium constant increases.
(D) The value of the equilibrium constant decreases.
(E) Neither the rate nor the value of the equilibrium constant is changed.

(A) (B) (C) (D) (E)

19. The critical temperature of a substance is the

- (A) temperature at which the vapor pressure of the liquid is equal to the external pressure
(B) temperature at which the vapor pressure of the liquid is equal to 760 mm Hg
(C) temperature at which the solid, liquid, and vapor phases are all in equilibrium
(D) temperature at which liquid and vapor phases are in equilibrium at 1 atmosphere
(E) lowest temperature above which a substance cannot be liquefied at any applied pressure

(A) (B) (C) (D) (E)



If the equilibrium constant for the reaction above is 3.7×10^{15} , which of the following correctly describes the standard voltage, E° , and the standard free energy change, ΔG° , for this reaction?

- (A) E° is positive and ΔG° is negative.
 (B) E° is negative and ΔG° is positive.
 (C) E° and ΔG° are both positive.
 (D) E° and ΔG° are both negative.
 (E) E° and ΔG° are both zero.

(A) (B) (C) (D) (E)

21. When $^{214}_{84}\text{Po}$ decays, the emission consists consecutively of an α particle, then two β particles, and finally another α particle. The resulting stable nucleus is

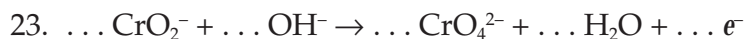
- (A) $^{206}_{83}\text{Bi}$ (B) $^{210}_{83}\text{Bi}$ (C) $^{206}_{82}\text{Bi}$
 (D) $^{208}_{82}\text{Pb}$ (E) $^{210}_{81}\text{Tl}$

(A) (B) (C) (D) (E)

22. The pH of 0.1-molar ammonia is approximately

- (A) 1 (B) 4 (C) 7 (D) 11 (E) 14

(A) (B) (C) (D) (E)



When the equation for the half-reaction above is balanced, what is the ratio of the coefficients $\text{OH}^- : \text{CrO}_2^-$?

- (A) 1 : 1 (B) 2 : 1 (C) 3 : 1
 (D) 4 : 1 (E) 5 : 1

(A) (B) (C) (D) (E)

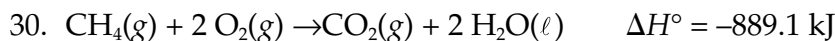


When the substances in the equation above are at equilibrium at pressure P and temperature T , the equilibrium can be shifted to favor the products by

- (A) increasing the pressure by means of a moving piston at constant T
 (B) increasing the pressure by adding an inert gas such as nitrogen
 (C) decreasing the temperature
 (D) allowing some gases to escape at constant P and T
 (E) adding a catalyst

(A) (B) (C) (D) (E)

25. The molality of the glucose in a 1.0-molar glucose solution can be obtained by using which of the following?
- (A) Volume of the solution
(B) Temperature of the solution
(C) Solubility of glucose in water
(D) Degree of dissociation of glucose
(E) Density of the solution
- (A) (B) (C) (D) (E)
26. The geometry of the SO_3 molecule is best described as
- (A) trigonal planar
(B) trigonal pyramidal
(C) square pyramidal
(D) bent
(E) tetrahedral
- (A) (B) (C) (D) (E)
27. Which of the following molecules has the shortest bond length?
- (A) N_2 (B) O_2 (C) Cl_2
(D) Br_2 (E) I_2
- (A) (B) (C) (D) (E)
28. What number of moles of O_2 is needed to produce 14.2 grams of P_4O_{10} from P? (Molecular mass $\text{P}_4\text{O}_{10} = 284$)
- (A) 0.0500 mole
(B) 0.0625 mole
(C) 0.125 mole
(D) 0.250 mole
(E) 0.500 mole
- (A) (B) (C) (D) (E)
29. If 0.060 faraday is passed through an electrolytic cell containing a solution of In^{3+} ions, the maximum number of moles of In that could be deposited at the cathode is
- (A) 0.010 mole
(B) 0.020 mole
(C) 0.030 mole
(D) 0.060 mole
(E) 0.18 mole
- (A) (B) (C) (D) (E)



$$\Delta H_f^\circ \text{H}_2\text{O}(\ell) = -285.8 \text{ kJ/mole}$$

$$\Delta H_f^\circ \text{CO}_2(\text{g}) = -393.3 \text{ kJ/mole}$$

What is the standard heat of formation of methane, $\Delta H_f^\circ \text{CH}_4(\text{g})$, as calculated from the data above?

- (A) -210.0 kJ/mole
 (B) -107.5 kJ/mole
 (C) -75.8 kJ/mole
 (D) 75.8 kJ/mole
 (E) 210.0 kJ/mole

(A) (B) (C) (D) (E)

31. Each of the following can act as both a Brønsted acid and a Brønsted base EXCEPT

- (A) HCO_3^- (B) H_2PO_4^- (C) NH_4^+
 (D) H_2O (E) HS^-

(A) (B) (C) (D) (E)

32. Two flexible containers for gases are at the same temperature and pressure. One holds 0.50 gram of hydrogen and the other holds 8.0 grams of oxygen. Which of the following statements regarding these gas samples is FALSE?

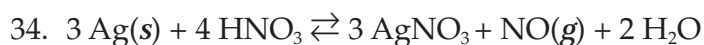
- (A) The volume of the hydrogen container is the same as the volume of the oxygen container.
 (B) The number of molecules in the hydrogen container is the same as the number of molecules in the oxygen container.
 (C) The density of the hydrogen sample is less than that of the oxygen sample.
 (D) The average kinetic energy of the hydrogen molecules is the same as the average kinetic energy of the oxygen molecules.
 (E) The average speed of the hydrogen molecules is the same as the average speed of the oxygen molecules.

(A) (B) (C) (D) (E)

33. Pi (π) bonding occurs in each of the following species EXCEPT

- (A) CO_2 (B) C_2H_4 (C) CN^-
 (D) C_6H_6 (E) CH_4

(A) (B) (C) (D) (E)



The reaction of silver metal and dilute nitric acid proceeds according to the equation above. If 0.10 mole of powdered silver is added to 10. milliliters of 6.0-molar nitric acid, the number of moles of NO gas that can be formed is

- (A) 0.015 mole
- (B) 0.020 mole
- (C) 0.030 mole
- (D) 0.045 mole
- (E) 0.090 mole

(A) (B) (C) (D) (E)

35. Which, if any, of the following species is in the greatest concentration in a 0.100-molar solution of H_2SO_4 in water?

- (A) H_2SO_4 molecules
- (B) H_3O^+ ions
- (C) HSO_4^- ions
- (D) SO_4^{2-} ions
- (E) All species are in equilibrium and therefore have the same concentrations.

(A) (B) (C) (D) (E)

36. At $20.^\circ\text{C}$, the vapor pressure of toluene is 22 mm Hg and that of benzene is 75 mm Hg. An ideal solution, equimolar in toluene and benzene, is prepared. At $20.^\circ\text{C}$, what is the mole fraction of benzene in the vapor in equilibrium with this solution?

- (A) 0.23 (B) 0.29 (C) 0.50
- (D) 0.77 (E) 0.83

(A) (B) (C) (D) (E)

37. Which of the following aqueous solutions has the highest boiling point?

- (A) 0.10 *M* potassium sulfate, K_2SO_4
- (B) 0.10 *M* hydrochloric acid, HCl
- (C) 0.10 *M* ammonium nitrate, NH_4NO_3
- (D) 0.10 *M* magnesium sulfate, MgSO_4
- (E) 0.20 *M* sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

(A) (B) (C) (D) (E)

38. When 70. milliliters of 3.0-molar Na_2CO_3 is added to 30. milliliters of 1.0-molar NaHCO_3 , the resulting concentration of Na^+ is
(A) 2.0 *M* (B) 2.4 *M* (C) 4.0 *M*
(D) 4.5 *M* (E) 7.0 *M* (A) (B) (C) (D) (E)
39. Which of the following species CANNOT function as an oxidizing agent?
(A) $\text{Cr}_2\text{O}_7^{2-}$ (B) MnO_4^- (C) NO_3^-
(D) S (E) I^- (A) (B) (C) (D) (E)
40. A student wishes to prepare 2.00 liters of 0.100-molar KIO_3 (molecular mass 214). The proper procedure is to weigh out
(A) 42.8 grams of KIO_3 and add 2.00 kilograms of H_2O
(B) 42.8 grams of KIO_3 and add H_2O until the final homogeneous solution has a volume of 2.00 liters
(C) 21.4 grams of KIO_3 and add H_2O until the final homogeneous solution has a volume of 2.00 liters
(D) 42.8 grams of KIO_3 and add 2.00 liters of H_2O
(E) 21.4 grams of KIO_3 and add 2.00 liters of H_2O (A) (B) (C) (D) (E)
41. A 20.0-milliliter sample of 0.200-molar K_2CO_3 solution is added to 30.0 milliliters of 0.400-molar $\text{Ba}(\text{NO}_3)_2$ solution. Barium carbonate precipitates. The concentration of barium ion, Ba^{2+} , in solution after reaction is
(A) 0.150 *M*
(B) 0.160 *M*
(C) 0.200 *M*
(D) 0.240 *M*
(E) 0.267 *M* (A) (B) (C) (D) (E)

42. One of the outermost electrons in a strontium atom in the ground state can be described by which of the following sets of four quantum numbers?
- (A) $5, 2, 0, \frac{1}{2}$
(B) $5, 1, 1, \frac{1}{2}$
(C) $5, 1, 0, \frac{1}{2}$
(D) $5, 0, 1, \frac{1}{2}$
(E) $5, 0, 0, \frac{1}{2}$ (A) (B) (C) (D) (E)
43. Which of the following reactions does NOT proceed significantly to the right in aqueous solutions?
- (A) $\text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2 \text{H}_2\text{O}$
(B) $\text{HCN} + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{CN}^-$
(C) $\text{Cu}(\text{H}_2\text{O})_4^{2+} + 4 \text{NH}_3 \rightarrow \text{Cu}(\text{NH}_3)_4^{2+} + 4 \text{H}_2\text{O}$
(D) $\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$
(E) $\text{H}_2\text{O} + \text{HSO}_4^- \rightarrow \text{H}_2\text{SO}_4 + \text{OH}^-$ (A) (B) (C) (D) (E)
44. A compound is heated to produce a gas whose molecular mass is to be determined. The gas is collected by displacing water in a water-filled flask inverted in a trough of water. Which of the following is necessary to calculate the molecular mass of the gas, but does NOT need to be measured during the experiment?
- (A) Mass of the compound used in the experiment
(B) Temperature of the water in the trough
(C) Vapor pressure of the water
(D) Barometric pressure
(E) Volume of water displaced from the flask (A) (B) (C) (D) (E)
45. A 27.0-gram sample of an unknown hydrocarbon was burned in excess oxygen to form 88.0 grams of carbon dioxide and 27.0 grams of water. What is a possible molecular formula of the hydrocarbon?
- (A) CH_4 (B) C_2H_2 (C) C_4H_3
(D) C_4H_6 (E) C_4H_{10} (A) (B) (C) (D) (E)

46. If the acid dissociation constant, K_a , for an acid HA is 8×10^{-4} at 25°C , what percent of the acid is dissociated in a 0.50-molar solution of HA at 25°C ?

- (A) 0.08% (B) 0.2% (C) 1%
(D) 2% (E) 4%

(A) (B) (C) (D) (E)

Study Resources

Textbooks that are frequently used in first-year chemistry courses can help you to prepare for the General Chemistry exam. You can find college-level general chemistry textbooks at most college bookstores. When selecting a textbook, check the table of contents against the “Knowledge and Skills Required” section on pages 1-5.

Since textbooks vary in their approach and emphasis, you should consult more than one textbook.

Additional suggestions for preparing for CLEP exams appear in Chapter 1.

Answers to Sample Questions

General Chemistry

- | | |
|-------|-------|
| 1. B | 26. A |
| 2. E | 27. A |
| 3. B | 28. D |
| 4. E | 29. B |
| 5. D | 30. C |
| 6. C | 31. C |
| 7. C | 32. E |
| 8. A | 33. E |
| 9. B | 34. A |
| 10. C | 35. B |
| 11. D | 36. D |
| 12. D | 37. A |
| 13. C | 38. D |
| 14. E | 39. E |
| 15. E | 40. B |
| 16. B | 41. B |
| 17. A | 42. E |
| 18. A | 43. E |
| 19. E | 44. C |
| 20. A | 45. D |
| 21. C | 46. E |
| 22. D | |
| 23. D | |
| 24. C | |
| 25. E | |
-